

TOPIC 18 — HL ACIDS/BASES

18.2 — BUFFER SOLUTIONS

IB Chemistry
T18D05



18.2 Buffer solutions - 2 hours

- 18.2.1 Describe the composition of a buffer solution and explain its action. (3)
- 18.2.2 Solve problems involving the composition and pH of a specified buffer system. (3)



18.2.1 – Definition of a Buffer Soln

- 18.2.1 Describe the composition of a buffer solution and explain its action. (3)
- A **buffer solution** is an aqueous solution whose pH (and $[H^+]$) remains unchanged by dilution with water or when relatively small amounts of acid or base are added to it. **Buffers resist change in pH.**
- Many processes require pH to be controlled such as enzymes working as catalysts, developing photographs, processes in soil, etc



18.2 – Types of Buffers

- There are three types of buffers:
 - Acidic buffers, which are prepared from a weak acid and a salt of the acid
 - For example, ethanoic acid and sodium ethanoate
 - Basic or alkaline buffers, which are prepared from a weak base and a salt of the base
 - For example, ammonia and ammonium chloride
 - Neutral buffers, which are prepared from phosphoric acid and its salts



18.2 – Action of a Buffer - Acids

- For our example of ethanoic acid and its salt, sodium ethanoate, the following reactions show few H^+ ions are in solution, but plenty of $\text{CH}_3\text{COO}^-(\text{aq})$
 - $\text{CH}_3\text{COONa}(\text{aq}) \rightarrow \text{CH}_3\text{COO}^-(\text{aq}) + \text{Na}^+(\text{aq})$
 - $\text{CH}_3\text{COOH}(\text{aq}) \rightleftharpoons \text{CH}_3\text{COO}^-(\text{aq}) + \text{H}^+(\text{aq})$
- If acid is added to the buffer, the $\text{CH}_3\text{COO}^-(\text{aq})$ above will “mop up” the H^+ ions producing more $\text{CH}_3\text{COOH}(\text{aq})$ as its equilibrium lies to the left
- If an alkali is added, the $\text{H}^+(\text{aq})$ will combine to form water and push the reaction to the left as needed to produce more H^+ for equilibrium
 - $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l})$



18.2 – Action of a Buffer - Bases

- For our example of ammonia and its salt, ammonium chloride, the following reactions show few OH^- ions are in solution, but plenty $\text{NH}_4^+(\text{aq})$
 - $\text{NH}_4\text{Cl}(\text{aq}) \rightarrow \text{NH}_4^+(\text{aq}) + \text{Cl}^-(\text{aq})$
 - $\text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$
- If alkali is added, the excess hydroxide will react with $\text{NH}_4^+(\text{aq})$ to form $\text{NH}_3(\text{aq})$ and $\text{H}_2\text{O}(\text{l})$ as the equilibrium is so far to the left, it “mops” up ions
- If an acid is added, the hydrogen ions combine with $\text{OH}^-(\text{aq})$ to form water:



18.2 – Calculations involving Buffers

- Consider the equilibrium for a weak acid:



- $$K_a = \frac{[\quad (\quad)]}{(\quad)}$$

- If we rearrange:

- $$[\text{H}^+(\text{aq})] = \frac{(\quad)}{(\quad)}$$

- Taking the $(-)\log_{10}$ of both sides:

- $$\text{pH} = \text{p}K_a - \log_{10} \frac{(\quad)}{(\quad)}$$

- $$\text{pH} = \text{p}K_a + \log_{10} \frac{(\quad)}{(\quad)}$$

This equation is known as the **Henderson-Hasselbalch equation**



18.2 – Henderson-Hasselbalch equation

- The **Henderson-Hasselbalch** equation indicates:
 - The pH of a buffer solution depends on the K_a of the weak acid
 - The pH of a buffer solution depends upon the ratio of the concentrations of the acid and its conjugate base and not their actual concentrations

$$\text{pH} = \text{pK}_a - \log_{10} \frac{(\quad)}{(\quad)}$$

$$\text{pH} = \text{pK}_a + \log_{10} \frac{(\quad)}{(\quad)}$$



18.2 – Characteristics of a Buffer

- **Dilution:** The H-H equation explains that the pH of a buffer will depend only on the [acid] and [conj base], so that dilution of the buffer should have no effect. If distilled water is added to a buffer solution, you dilute both components of the buffer equally.
- **Buffering Capacity:** explains the ability of the buffer to resist change in pH. The bufferin capacity increases as the [buffer salt/acid] solution increases. The closer the pH is to the pK_a , the greater the buffer capacity. Expressed as [NaOH] required to increase pH by 1.0



18.2 – Calc #1 – pH of Buffer

- Calculate the pH of a buffer containing 0.20 moles of sodium ethanoate in 500cm³ of 0.10 mol dm⁻³ ethanoic acid. K_a for ethanoic acid is 1.8×10^{-5}
 - $\text{CH}_3\text{COONa}(\text{aq}) \rightarrow \text{CH}_3\text{COO}^-(\text{aq}) + \text{Na}^+(\text{aq})$
 - $\text{CH}_3\text{COOH}(\text{aq}) \rightleftharpoons \text{CH}_3\text{COO}^-(\text{aq}) + \text{H}^+(\text{aq})$
- $[\text{CH}_3\text{COO}^-(\text{aq})] = \frac{\quad}{\quad} = 0.40 \text{ mol dm}^{-3}$
- $K_a = \frac{[\quad]}{[\quad]} = 1.8 \times 10^{-5}$
- $1.8 \times 10^{-5} = \frac{(\quad)}{\quad}$
 - $x = 4.5 \times 10^{-6} = [\text{H}^+(\text{aq})]$
 - $\text{pH} = -\log_{10}[\text{H}^+] = 5.3$



18.2 – Calc #2 – Mass of Req'd Salt

- Calculate the mass of sodium propanoate ($M=96.07\text{g mol}^{-1}$) that must be dissolved in 1.00dm^3 of propanoic acid ($\text{pK}_a = 4.87$) to give a buffer solution with pH of 4.5.
 - Let x represent the [propanoate ions] and y represent amount of sodium propanoate
 - $[\text{H}^+] = 10^{-\text{pH}} = 10^{-4.5} = 3.16 \times 10^{-5} \text{ mol dm}^{-3}$
 - $K_a = 10^{-\text{pK}_a} = 10^{-4.87} = 1.35 \times 10^{-5}$
 - $K_a = \frac{[\text{ } (\text{ })]}{(\text{ })} = 1.35 \times 10^{-5}$
 - $1.35 \times 10^{-5} = \text{_____}$



$$x = 0.427 \text{ mol dm}^{-3} = \text{_____}$$

$$\text{mol} = 0.427$$

- $96.07\text{g mol}^{-1} \times 0.427\text{mol} = \mathbf{41.0\text{g}}$

18.2 – Calc #3 – pH of Buffer after base is added

- A buffer contains 0.20mol of sodium ethanoate in 500cm³ of 0.10 mol dm⁻³ ethanoic acid. K_a for ethanoic acid is 1.8×10^{-5} : Calculate the pH after 0.025 moles of sodium hydroxide is added

- *The addition of OH⁻ ions will cause the dissociation to shift to the right making more H⁺ ions available. So the amount of OH⁻ ions must be subtracted from the ethanoic acid and added to the amount of ethanoate ions. $\text{CH}_3\text{COOH}(\text{aq}) \rightleftharpoons \text{CH}_3\text{COO}^-(\text{aq}) + \text{H}^+(\text{aq})$*

- $K_a = \frac{[\quad (\quad)]}{(\quad)} = 1.8 \times 10^{-5}$

- $[\text{H}^+] = x$

- $[\text{CH}_3\text{COO}^-(\text{aq})] = \text{—————} = 0.45 \text{ mol dm}^{-3}$

$$[\text{CH}_3\text{COOH}(\text{aq})] = (0.10 \text{ mol dm}^{-3} - \text{—————}) - x = 0.050 \text{ mol dm}^{-3} - x$$

Continued on next slide....



18.2 – Calc #3 Continued

- $[\text{CH}_3\text{COOH}(\text{aq})] = 0.050 \text{ mol dm}^{-3} - x \approx 0.050 \text{ mol dm}^{-3}$ (x is small)
- $K_a = 1.8 \times 10^{-5} = \frac{[\quad]}{[\quad]} = \quad - x \approx \quad$
- $x = 2.0 \times 10^{-6} = [\text{H}^+]$
- $\text{pH} = -\log_{10}[\text{H}^+] = -\log[2.0 \times 10^{-6}] = 5.7$

